Chem 1515
Problem Set #11
Fall 2001

Name______________________________

TA Name ________________________

Lab Section #______

ALL work must be shown to receive full credit. **Due at the beginning of lecture on Friday, November 2, 2001.**

PS11.1. For aqueous solutions of the following substances, write the dissociation reaction and indicate whether the substance behaves as an Arrhenius acid or base.

a) $HF(aq) \rightleftarrows \mathbf{H^+(aq) + F^-(aq)}$ **acid**

b) $HC_6H_5O(aq) \rightleftarrows \mathbf{H^+(aq) + C_6H_5O^-(aq)}$ **acid**

c) $Ba(OH)_2(aq) \rightarrow \mathbf{Ba^{2+}(aq) + 2OH^-(aq)}$ **base**

d) $LiOH(aq) \rightarrow \mathbf{Li^+(aq) + OH^-(aq)}$ **base**

e) $H_2O(aq) \rightleftarrows \mathbf{H^+(aq) + OH^-(aq)}$ **neutral**

f) $H_2CO_3(aq) \rightleftarrows \mathbf{H^+(aq) + HCO_3^-(aq)}$ **acid**

PS11.2. Calculate the pH and pOH in each of the following aqueous solutions. In each case, indicate whether the solution is acidic or basic.

a) $[H^+] = 3.89x\ 10^{-5}$ M
pH = 4.41
pOH = 9.59
acidic

b) $[OH^-] = 8.34 \times 10^{-2}$ M
pH = 12.92
pOH = 1.08 basic

c) $[OH^-] = = 1.50 \times 10^{-7}$ M ($[OH^-]$ in milk)
pH = 7.18
pOH = 6.82 basic

d) $[H^+] = 9.39 \times 10^{-10}$ M
pH = 9.02
pOH = 4.97 basic

e) $[H^+] = 4.0$ M
pH = -0.60
pOH = 14.6 acidic

f) $[OH^-] = 10.1$ M
pH = 15.0
pOH = -1.00
basic

PS11.3. Calculate the $[H^+]$ and $[OH^-]$ in each of the following aqueous solutions.

$$\mathbf{pH = -log[H^+]}$$

$$\mathbf{-3.40 = log[H^+]}$$

$$\mathbf{10^{-3.40} = 10^{log[H^+]}}$$

$$\mathbf{3.98 \times 10^{-4}\ M = [H^+]}$$

$$\mathbf{K_w = 1 \times 10^{-14} = [H^+][OH^-]}$$

$$\mathbf{[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}}$$

$$\mathbf{[OH^-] = \frac{1 \times 10^{-14}}{3.98 \times 10^{-4}\ M}}$$

$$\mathbf{[OH^-] = 2.51 \times 10^{-11}\ M}$$

b) pH = 6.7 (pH of salvia)
$\mathbf{[H^+] = 1.99 \times 10^{-7}\ M}$
$\mathbf{[OH^-] = 5.01 \times 10^{-8}\ M}$

$\mathbf{[H^+] = 3.98 \times 10^{-5}\ M}$
$\mathbf{[OH^-] = 2.51 \times 10^{-10}\ M}$

d) pOH = 2.15
$\mathbf{[H^+] = 1.41 \times 10^{-12}\ M}$
$\mathbf{[OH^-] = 7.07 \times 10^{-3}\ M}$

e) pOH = 12.4
$\mathbf{[H^+] = 2.51 \times 10^{-2}\ M}$
$\mathbf{[OH^-] = 3.98 \times 10^{-13}\ M}$

f) pH = -0.650
$\mathbf{[H^+] = 4.47\ M}$
$\mathbf{[OH^-] = 2.24 \times 10^{-15}\ M}$

PS11.4. For each of the following acids, write the formula for the conjugate base.

a) HPO_4^{2-} **PO_4^{3-}**
c) H_2O **OH^-**
e) OH^- **O^{2-}**
b) $HClO_3$ **ClO_3^-**
d) $CH_3CH_2NH_3^+$ **$CH_3CH_2NH_2$**
f) NH_4^+ **NH_3**

PS11.5. For each of the following bases, write the formula for the conjugate acid.

a) OH^- **H_2O**
c) HCO_3^{2-} **H_2CO_3**
e) CH_3NH_2 **$CH_3NH_3^+$**
b) Cl^- **HCl**
d) H_2O **H_3O^+**
f) $(CH_3)_3N$ **$(CH_3)_3NH^+$**

PS11.6. For the following compounds, write the reaction with water and indicate the Brø nsted acid, base, the conjugate acid and conjugate base.

a) $HBr(g) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{Br^-}(aq)$
A B CA CB

b) $NH_3(g) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{NH_4^+}(aq) + \mathbf{OH^-}(aq)$
B A CA CB

c) $HCN(g) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{CN^-}(aq)$
A B CA CB

d) $HC_7H_5O_2(s) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{C_7H_5O_2^-}(aq)$
A B CA CB

e) $CH_3NH_2(l) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{CH_3NH_3^+}(aq) + \mathbf{OH^-}(aq)$
B A CA CB

PS11.7. For each of the following compounds, write two Brø nsted-Lowry equations, one showing how the substance behaves as an acid, the second showing how the substance behaves as a base.

a) $HCO_3^-(aq)$

$\mathbf{HCO_3^-}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{CO_3^{2-}}(aq)$ **acid**

$\mathbf{HCO_3^-}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{OH^-}(aq) + \mathbf{H_2CO_3}(aq)$ **base**

b) $NH_3(aq)$

$\mathbf{NH_3}(g) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{NH_4^+}(aq) + \mathbf{OH^-}(aq)$ **base**

$\mathbf{NH_3}(g) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{NH_2^-}(aq) + \mathbf{H_3O^+}(aq)$ **acid**

c) $HPO_4^{2-}(aq)$

$\mathbf{HPO_4^{2-}}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{PO_4^{3-}}(aq)$ **acid**

$\mathbf{HPO_4^{2-}}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{OH^-}(aq) + \mathbf{H_2PO_4^-}(aq)$ **base**

d) $HSO_4^-(aq)$

$\mathbf{HSO_4^-}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{H_3O^+}(aq) + \mathbf{SO_4^{2-}}(aq)$ **acid**

$\mathbf{HSO_4^-}(aq) + \mathbf{H_2O}(l) \rightleftarrows \mathbf{OH^-}(aq) + \mathbf{H_2SO_4}(aq)$ **base**

PS11.8. Determine the equilibrium constant for the following solutions. (Show your work clearly!)

a) 0.250 M HF whose pH = 1.89.

	HF	$\rightleftarrows$	H^+	+	F^-	
I	.250		~0		0	
C	-x		+x		+x	x = $[HF]_D$
E	.250 - x		x		x	

$$K_a = \frac{[H^+][F^-]}{[HF]}$$

Since the pH = 1.89, the $[H^+]$ = 1.29 x 10^{-2} M which is equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.250 - x)}$$

Substituting for x,

$$K_a = \frac{(1.29 \times 10^{-2})^2}{.250 - 0.0129} = 7.02 \times 10^{-4}$$

b) 0.235 M NH_3 whose pH = 11.31.

	NH_3	+ H_2O	$\rightleftarrows$	$NH_4^+(aq)$ +	$OH^-(aq)$	
I	0.235 M	-		0	0	
C	-x	-		+x	+x	let x = $[NH_3]_R$
E	0.235 - x	-		0 + x	0 + x	

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Since the pH = 11.31, the pOH = 2.69 and the $[OH^-]$ = 2.0 x 10^{-3} M which is also equal to x as shown in the ICE table above.

$$K_b = \frac{(x)(x)}{(0.1 - x)} = \frac{(2.0 \times 10^{-3})(2.0 \times 10^{-3})}{(0.235 - 2.0 \times 10^{-3})} = 1.79 \times 10^{-5}$$

c) 0.500 M B whose pH = 9.34.

	B	+ H_2O	$\rightleftarrows$	$BH^+(aq)$ +	$OH^-(aq)$	
I	0.500 M	-		0	0	
C	-x	-		+x	+x	let x = $[B]_R$
E	0.500 - x	-		0 + x	0 + x	

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Since the pH = 9.34, the pOH = 4.66 and the $[OH^-]$ = 2.19 x 10^{-5} M which is also equal to x as shown in the ICE table above.

$$K_b = \frac{(x)(x)}{(0.5 - x)}$$

Substituting for x,

$$= \frac{(2.19 \times 10^{-5})(2.19 \times 10^{-5})}{(0.500 - 2.19 \times 10^{-5})}$$

$$K_b = \frac{(2.19 \times 10^{-5})^2}{0.500} = 9.57 \times 10^{-10}$$

d) 0.302 M HA whose pH = 4.80.

	HA	$\rightleftarrows$	H^+	+	A^-	
I	0.302		~0		0	
C	-x		+x		+x	x = $[HA]_D$
E	0.302 - x		x		x	

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Since the pH = 4.80, the $[H^+] = 1.58 \times 10^{-5}$ M which is also equal to x as shown in the ICE table above.

$$K_a = \frac{(x)(x)}{(0.302 - x)}$$

Substituting for x,

$$K_a = \frac{(1.58 \times 10^{-5})^2}{.302 - 1.58 \times 10^{-5}}$$

$$= \frac{(1.58 \times 10^{-5})^2}{.302} = 8.32 \times 10^{-10}$$

PS11.9. Given the following substances and their initial concentration:

a)	0.200 M HNO_3	e)	55.5 M H_2O	i)	0.200 M HC_6H_5O
b)	0.200 M HF	f)	0.200 M HNO_2	j)	0.200 M $Ba(OH)_2$
c)	0.200 M NaOH	g)	0.200 M CH_3NH_2	k)	0.003501 M HF
d)	0.200 M C_5H_5N	h)	0.200 M $C_2H_5NH_2$	l)	0.200 M HOCl

Answer the following,

i) identify each as an acid, base or neutral substance.

a)	0.200 M HNO_3	**A**	e)	55.5 M H_2O	**N**	i)	0.200 M HC_6H_5O	**A**
b)	0.200 M HF	**A**	f)	0.200 M HNO_2	**A**	j)	0.200 M $Ba(OH)_2$	**B**
c)	0.200 M NaOH	**B**	g)	0.200 M CH_3NH_2	**B**	k)	0.003501 M HF	**A**
d)	0.200 M C_5H_5N	**B**	h)	0.200 M $C_2H_5NH_2$	**B**	l)	0.200 M HOCl	**A**

ii) list the K_a value for each acid and K_b value for each base.

a) 0.200 M HNO_3	**K_a is very large**
b) 0.200 M HF	**$K_a = 6.8 \times 10^{-4}$**
c) 0.200 M NaOH	**K_b is very large**
d) 0.200 M C_5H_5N	**$K_b = 1.7 \times 10^{-9}$**
e) 55.5 M H_2O	**$K_a = 1 \times 10^{-14}$**
f) 0.200 M HNO_2	**$K_a = 4.5 \times 10^{-4}$**
g) 0.200 M CH_3NH_2	**$K_b = 4.4 \times 10^{-4}$**
h) 0.200 M $C_2H_5NH_2$	**$K_b = 6.4 \times 10^{-4}$**
i) 0.200 M HC_6H_5O	**$K_a = 1.3 \times 10^{-10}$**
j) 0.200 M $Ba(OH)_2$	**K_b is very large**
k) 0.00491 M HF	**$K_a = 6.8 \times 10^{-4}$**
l) 0.200 M HOCl	**$K_b = 3.0 \times 10^{-8}$**

iii) identify each substance as strong (S) or weak (W).

a)	0.200 M HNO_3	**SA**	e)	55.5 M H_2O	**N**	i)	0.200 M HC_6H_5O	**WA**
b)	0.200 M HF	**WA**	f)	0.200 M HNO_2	**WA**	j)	0.200 M $Ba(OH)_2$	**SB**
c)	0.200 M NaOH	**SB**	g)	0.200 M CH_3NH_2	**WB**	k)	0.003501 M HF	**WA**
d)	0.200 M C_5H_5N	**WB**	h)	0.200 M $C_2H_5NH_2$	**WB**	l)	0.200 M HOCl	**WA**

iv) calculate the $[H^+]$ and the pH of each of the solutions.

a) 0.200 M HNO_3

	$HNO_3 \rightarrow$	H^+	+	NO_3^-
I	**.2**		**~0**	**0**
F	**0**		**.2**	**.2**

HNO_3 is a strong acid and completely dissociates in water.

Therefore, $[H^+] = 0.2$ M and pH = 0.700

b) 0.200 M HF **$[H^+] = 1.2 \times 10^{-2}$ M pH = 1.93**

c) 0.100 M NaOH

	$NaOH \rightarrow$	Na^+	+	OH^-
I	**.2**		**0**	**~0**
F	**0**		**.2**	**.2**

NaOH is a strong base and completely dissociates in water. Therefore,

$[OH^-] = 0.2$ M and pOH = 0.70, pH = 13.3: $[H^+] = 5.0 \times 10^{-14}$ M

d) 0.200 M C_5H_5N

	C_5H_5N	$+ H_2O$	$\rightleftarrows$	$C_5H_5NH^+(aq)$	$+ OH^-(aq)$	
I	**0.2 M**	**-**		**0**	**0**	
C	**-x**	**-**		**+x**	**+x**	**let x = $[C_5H_5N]_R$**
E	**0.2 - x**	**-**		**0+x**	**0+x**	

$$Kb = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$$

$$1.7 \times 10^{-9} = \frac{(x)\ (x)}{0.2 - x} \quad \text{assume } x << 0.2$$

$$1.7 \times 10^{-9} = \frac{x^2}{.2}$$

$$3.4 \times 10^{-10} = x^2$$

$$1.8 \times 10^{-5}\ M = x = [OH^-]$$

pOH = 4.73

pH = 9.27: $[H^+] = 5.4 \times 10^{-10}$ M

e) 55.5 M H_2O **$[H^+] = 1.0 \times 10^{-7}$ M: pH = 7.00**

f) 0.200 M HNO_2 **$[H^+] = 9.5 \times 10^{-3}$ M: pH = 2.02**

g) 0.200 M CH_3NH_2 **$[OH^-] = 9.4 \times 10^{-3}$ M and pOH = 2.03, pH = 11.98: $[H^+] = 1.07 \times 10^{-12}$ M**

h) h) 0.200 M $C_2H_5NH_2$

	$C_2H_5NH_2$	$+ H_2O$	$\rightleftarrows$	$C_2H_5NH_3^+(aq)$	$+ OH^-(aq)$	
I	**0.2 M**	**-**		**0**	**0**	
C	**-x**	**-**		**+x**	**+x**	**let x = $[C_2H_5NH_2]_R$**
E	**0.2 - x**	**-**		**0+x**	**0+x**	

$$Kb = \frac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$$

$$6.4 \times 10^{-4} = \frac{(x)\ (x)}{0.2 - x} \quad \text{assume } x << 0.1$$

$$6.4 \times 10^{-4} = \frac{x^2}{.2}$$

$$1.28 \times 10^{-4} = x^2$$

$$1.1 \times 10^{-2}\ M = x = [OH^-]$$

pOH = 1.96

pH = 12.04: $[H^+] = 9.09 \times 10^{-13}$

i) 0.200 M HC_6H_5O **$[H^+] = 5.1 \times 10^{-6}$ M: pH = 5.30**

j) 0.100 M $Ba(OH)_2$

	$Ba(OH)_2$	$\rightarrow$	Ba^{2+} +	$2OH^-$
I	**.2**		**0**	**~0**
F	**0**		**.2**	**.4**

$Ba(OH)_2$ is a strong base and completely dissociates in water. Therefore,

$[OH^-] = 0.4$ M and pOH = 0.40, pH = 13.6: $[H^+] = 2.5 \times 10^{-14}$ M

k) 0.00350 M HF

	HF	⇄	H^+ + F^-		
I	**.00350**		**~0**	**0**	
C	**-x**		**+x**	**+x**	$x = [HF]_{diss}$
E	**.00350-x**		**x**	**x**	

$$Ka = \frac{[H+][F^-]}{[HF]}$$

$$6.8 \times 10^{-4} = \frac{(x)(x)}{.00350-x}$$

$$6.8 \times 10^{-4} = \frac{x^2}{.00350-x}$$

$$2.38 \times 10^{-6} - 6.8 \times 10^{-4}x = x^2$$

$$2.38 \times 10^{-6} - 6.8 \times 10^{-4}x - x^2 = 0$$

$$x^2 + 6.8 \times 10^{-4}x - 2.38 \times 10^{-6} = 0$$

solving the quadratic equation $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$$x = \frac{-6.8 \times 10^{-4} \pm \sqrt{(6.8 \times 10^{-4})^2 - 4(1)(-2.38 \times 10^{-6})}}{2(1)}$$

$$x = \frac{-6.8 \times 10^{-4} \pm 3.16 \times 10^{-3}}{2}$$

Use only the positive root

$x = 1.24 \times 10^{-3}$ **M** $= [H^+]$
pH = 2.91

l) 0.200 M HOCl **$[H^+] = 7.5 \times 10^{-5}$ M: pH = 4.11**

v) determine the percent ionization for each acid and base.
w)

a)	0.200 M HNO_3	**100%**
b)	0.200 M HF	**5.9%**
c)	0.200 M NaOH	**100%**
d)	0.200 M C_5H_5N	**< 1%**
e)	55.5 M H_2O	**< 1%**
f)	0.200 M HNO_2	**4.8%**
g)	0.200 M CH_3NH_2	**4.7%**
h)	0.200 M $C_2H_5NH_2$	**5.5%**
i)	0.200 M HC_6H_5O	**< 1%**
j)	0.200 M $Ba(OH)_2$	**100%**
k)	0.003501 M HF	**35.1%**
l)	0.200 M HOCl	**< 1%**

vi) rank all substances from strongest acid...weakest acid...neutrals.. ...weakest base...strongest base.

a) 0.200 M HNO_3	**pH = 0.700**
b) 0.200 M HF	**pH = 1.93**
f) 0.200 M HNO_2	**pH = 2.02**
l) 0.200 M HOCl	**pH = 4.11**
i) 0.200 M HC_6H_5O	**pH = 5.30**
e) 55.5 M H_2O	**pH = 7.00**
d) 0.200 M C_5H_5N	**pH = 9.27**
g) 0.200 M CH_3NH_2	**pH = 11.98**
h) 0.200 M $C_2H_5NH_2$	**pH = 12.04**
c) 0.200 M NaOH	**pH = 13.3**
j) 0.200 M $Ba(OH)_2$	**pH = 13.6**